IB Chemistry | Brakke | ECA – Topic 05

T05D01 HW: Thermochemistry II

Name KEY

Directions: ***Solve each of the following on a sheet of your own paper.*** Show ALL work including UNITS. If specific heats for some of the elements are needed and not given consult your periodic table.

1. A piece of silver has a mass of 362 grams and a heat capacity of 85.7 J/degree C. What is the specific heat of silver? [Think of the units associated with specific heat, and an epiphany will hopefully occur.]

$$\frac{85.7\ J}{c} \times \frac{1}{362\ g} = 0.237\ J/g°C = c_{Ag}$$

2. Calculate the heat liberated when 366 grams of mercury cools from 77.0 to 12.0 degrees C.

$q = ?$
$m = 366 g$
$c = 0.140\ J/g°C$
$\Delta T = 12 - 77 = -65°C$

$q = mc\Delta T$
$q = (366 g)(0.140\ J/g°C)(-65°C)$
$q = -3330.6\ J$ or $-3.33\ kJ$

3. A piece of copper of mass 6.22 kg is heated from 20.5 degrees C to 324.3 degrees C. Calculate the heat absorbed by the metal.

$q = ?$
$m = 6.22\ kg$ or $6{,}220\ g$
$c = 0.385\ J/g°C$
$\Delta T = 324.3 - 20.5 = 303.8°C$

$q = mc\Delta T$
$(6220\ g)(0.385\ J/g°C)(303.8°C)$
$q = 727510\ J \approx 728\ kJ$

4. A sheet of gold weighing 10.6 g and at a temperature of 18.0 degrees C is placed on a flat sheet of iron weighing 20.0 grams and at a temperature of 55.6 degrees C. What is the final temperature of the combined metals? Assume **NO** heat is lost to the surroundings.

$m_{Au} = 10.6 g$
$c_{Au} = 0.129\ J/g°C$
$\Delta T_{Au} = T_f - 18.0°C$

$m_{Fe} = 20.0 g$
$c_{Fe} = 0.449\ J/g°C$
$\Delta T_{Fe} = T_f - 55.6°C$

$(mc\Delta T)_{Au} = -(mc\Delta T)_{Fe}$

$\underbrace{(10.6 g)(0.129\ J/g°C)}_{1.37}(T_f - 18°C) = -\overbrace{(20.0 g)(0.449\ J/g°C)}^{8.8}(T_f - 55.6°C)$

$1.37T_f - 24.66 = -8.8T_f + 489.28$
$10.17T_f = 513.94$
$T_f = 50.53°C$

5. Calculate the initial temperature of 50.0 grams of cold water when 85.5 grams of metallic iron at 99.8 degrees C is placed into it. The final temperature of the mixture is 22.3 degrees C. Assume no heat loss to the surroundings.

$m_{H_2O} = 50.0 g$
$c_{H_2O} = 4.184\ J/g°C$
$\Delta T = 22.3 - T_i$

$m_{Fe} = 85.5\ g$
$c_{Fe} =$
$\Delta T = 22.3°C - 99.8°C = -77.5°C$

6. A certain mass of water was heated with 41.8 kJ, raising its temperature from 22.0 °C to 28.5 °C. Find the mass of water

$q = mc\Delta T$

$41.8\ kJ = 41{,}800\ J$

$41{,}800\ J = m_{H_2O}(4.184\ J/g°C)(6.5°C)$

$\boxed{m_{H_2O} = 1537\ g \text{ or } 1.537\ kg}$

7. When 80.0 grams of a certain metal at 90.0 °C was mixed with 100.0 grams of water at 30.0 °C, the final equilibrium temperature of the mixture was 36.0 °C. What is the specific heat of the metal?

$m_{met} = 80.0\ g$

$\Delta T_{met} = -44°C$

$c_{met} = ?$

$q_{met} = -q_{H_2O}$

$(mc\Delta T)_{met} = -(mc\Delta T)_{H_2O}$

$(80.0\ g)(c_{met})(-44°C) = -(100.0\ g)(4.184\ J/g°C)(6°C)$

$\boxed{c_{met} = 0.713\ J/g°C}$

$m_{H_2O} = 100\ g$

$\Delta T_{H_2O} = 6°C$

$c_{H_2O} = 4.184\ J/g°C$

8. 10.0 g of a fuel are burned under a calorimeter containing 200.0 g of H_2O. The temperature of the water increases from 15.0 °C to 55.0 °C. Calculate the total heat produced (in joules) and the heat of combustion per gram of fuel.

$m_{fuel} = 10.0\ g$

$q_{fuel} = \frac{-33{,}472\ J}{10.0\ g} = -3{,}347.2\ J/g$

$q_{H_2O} = mc\Delta T$

$q_{H_2O} = (200.0\ g)(4.184\ J/g°C)(40°C)$

$q_{H_2O} = 33{,}472\ J$

$q_{fuel} = -(q_{H_2O})$

$q_{fuel} = -33{,}472\ J$

$m_{H_2O} = 200.0\ g$

$\Delta T_{H_2O} = 40°C$

$c_{H_2O} = 4.184\ J/g°C$

9. When 12.29 g of finely divided brass (60% Cu, 40% Zn) at 95.0°C is quickly stirred into 40.00 g of water at 22.0°C in a calorimeter, the water temperature rises to 24.0°C. Find the specific heat of brass.

$m_{brass} = 12.29\ g$

$\Delta T_{brass} = -71°C$

$c_{brass} = ?$

$q_{brass} = -(q_{H_2O})$

$(mc\Delta T)_{brass} = -(mc\Delta T)_{H_2O}$

$(12.29\ g)(c_{brass})(-71.0°C) = -(40.00\ g)(4.184\ J/g°C)(2°C)$

$\boxed{c_{brass} = 0.384\ J/g°C}$

$m_{H_2O} = 40.00\ g$

$c_{H_2O} = 4.184\ J/g°C$

$\Delta T_{H_2O} = 2°C$

10. A 25.0 g piece of aluminum (which has a molar heat capacity of 24.03J/°C mol) is heated to 82.4°C and dropped into a calorimeter containing water (specific heat capacity of water is 4.18J/g°C) initially at 22.3°C. The final temperature of the water is 24.9°C. Calculate the mass of water in the calorimeter.

$m_{Al} = 25.0\ g$

$C_{Al} = 24.03\ J/°C\ mol$

$\Delta T_{Al} = -57.5°C$

$q_{Al} = -q_{H_2O}$

$(mc\Delta T)_{Al} = -(mc\Delta T)_{H_2O}$

$(25.0\ g)(0.891\ J/g°C)(-57.5°C) = -(m_{H_2O})(4.184\ J/g°C)(2.6°C)$

$\boxed{m_{H_2O} = 117.7\ g}$

$m_{H_2O} = ?$

$c_{H_2O} = 4.184\ J/g°C$

$\Delta T_{H_2O} = 2.6°C$

$\frac{24.03\ J}{°C\ mol} \times \frac{1\ mol\ Al}{26.982} = 0.891\ J/g°C$